**Special Notes:**

HA = acid

B = B-L base

BOH = Arr. Base

A- = conjugate base

BH+ = conjugate acid (for B-L Base)

B+ = conjugate acid (ONLY FOR Arr. Base)

Ka = equilibrium constant for an acid. MUST have H+ in the numerator of the equation

Kb = equilibrium constant for a base. MUST have OH- in the numerator of the equation

Kw = equilibrium constant for water. Kw = 10^-14 = K_a x K_b

pH = -log[H+]

pOH = -log[OH-]

pH + pOH = 14

pKa = -log[Ka]

pKb = -log[Kb]

Henderson-Hasselbach Equation for BUFFERS:

pH = pKa + log[A-]/[HA]

pOH = pKb + log[BH+]/B

Calculating the pH of a Solution Reference Guide

GT/AP Chemistry
Equilibrium & Acid-Base
Big Idea #6
**pH of a Strong Acid**

Strong acids dissociate completely and produce a proton (H⁺). The [H⁺] is equal to the [HA]. The pH can be calculated from the molarity of the acid directly.

List & Name the Six Strong Acids:

**Example:** HNO₃

Write the reaction for HNO₃ in water:

Draw a beaker diagram for HNO₃ in water.

Find the pH of a 0.01M HNO₃ solution.

What is the pH of a buffer that is made with 100mL of 0.01M HNO₃ and 30mL of 0.01M NaOH (molarity will change because the volume will change!!)?

What is the pH of a buffer that is made with 100mL of 0.01M CH₃CH₂NH₂ and 60mL of 0.01M HCl (Molarities will change because the volume will change!!)?
**pH Buffers**

In order to calculate the pH of a buffer, you must use the Henderson-Hasselbach Equation (see back).

What is the pH of a buffer that is made with 50mL of 0.01M HNO$_2$ and 60mL of 0.01M KNO$_2$ (The molarities will change because the volume will change!!)?

**pH of a Strong Base**

Strong bases dissociate completely and produce a hydroxide ion (OH$^-$. The [OH$^-$] is equal to the [BOH]. The pOH can be calculated from the Molarity of the base directly. Then the pOH can be converted to pH.

How will you be able to recognize a strong base?

**Example:** KOH

Write the reaction for KOH in water:

Draw a beaker diagram for KOH in water.

What is the pH of a buffer that is made with 50mL of 0.01M CH$_3$CH$_2$NH$_2$ and 40mL of 0.01M CH$_3$CH$_2$NH$_3$Cl (The molarities will change because the volume will change!!)?

Find the pH of a 0.01M KOH solution.
**pH of a Weak Acid (Monoprotic)**

Weak acids DO NOT dissociate completely. They produce some protons (H\(^+\)), but most stays in the HA form. These reactions are always at equilibrium. How much dissociates is based in the K\(_a\) value. The pH can be calculated by solving an ICE chart for the acid and using the K\(_a\) Value.

Find 3 inorganic weak acids (monoprotic) and their K\(_a\) values. Put in order from strongest to weakest (all are weak in general)

Find 2 organic weak acids (monoprotic) and their K\(_a\) values. Put in order from strongest to weakest (all are weak in general).

**Example:** HNO\(_2\)

Write the reaction for HNO\(_2\) in water. Label the conjugate base.

Draw a beaker diagram for HNO\(_2\) in water.
Buffers

Buffers are solutions which can resist changes in pH. The hold a maintain a relatively stable pH. Every buffer has a limit known as a buffer capacity which is related to the concentration of the buffer (the more concentrated, the greater the buffering capacity). In order to have a buffer, you must have a solution which contains both a weak acid and its conjugate base, or a weak base and its conjugate acid. In order to make a buffer you can add an acid and a salt containing a conjugate base. You can also form a conjugate base by starting with an acid and adding a strong base to it until some of the acid is turned into a conjugate base. You can also add a base and a salt containing its conjugate acid. Or you can start with a base and add strong acid until some of the base has been turned into the conjugate acid.

Describe two methods of creating a buffer for each of the following:

100mL of 0.01M HNO$_2$

Find the pH of a 0.01M HNO$_2$ solution.

Draw a beaker diagram of the buffer created:
**pH of a Weak Base**

Weak bases DO NOT ionize completely. They produce some hydroxide ions (OH\(^-\)), but most stays in the B form. These reactions are always at equilibrium. How much dissociates is based in the \(K_b\) value. The pOH can be calculated by solving an ICE chart for the base and using the \(K_b\) value. The pOH can then be converted to pH.

Find 3 weak bases. List their \(K_b\) values. Put them in order from strongest base to weakest base (all are weak generally)

Write the reactions for the following weak bases in water. Label the conjugate acid for each.

\[ \text{NH}_3 \]
\[ \text{CH}_3\text{NH}_2 \]
\[ (\text{CH}_3)_2\text{NH} \]

**Example:**
Write the reaction for \(\text{CH}_3\text{CH}_2\text{NH}_2\) in water. Label the conjugate base.

Draw a beaker diagram for \(\text{CH}_3\text{CH}_2\text{NH}_2\) in water.

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**Categorizing**

The most important step before calculating pH is determining WHAT YOU HAVE!

Label each one of the following as strong acid, weak acid, acidic salt, strong base, weak base, basic salt, neutral salt:

KF
HCl
HNO\(_2\)
NH\(_4\)Cl
CH\(_3\)CH\(_2\)CH\(_2\)NH\(_2\)
NaCN
KOH
KCl
CH\(_3\)COOH
(CH\(_3\))\(_2\)NH\(_2\)Br
NH\(_4\)NO\(_3\)
Determining pH Level of Salts

All salts contain a cation (positive charge) and anion (negative charge). To determine acidity of the salt, first break into two components. Cross out the neutral ions present. If the both are neutral, the salt is neutral. If the remaining ion is the cation, it is an acidic salt. If the remaining ion is an anion, it is a basic salt.

List the neutral cations:

List the neutral anions:

Examples:
Circle acidic salts, square basic salts. Underline neutral salts.

\[ \text{KNO}_3 \quad \text{LiClO}_4 \]
\[ \text{NaNO}_2 \quad (\text{CH}_3)_2\text{NHCl} \]
\[ \text{NH}_4\text{NO}_3 \quad \text{NaCN} \]
\[ \text{CH}_3\text{CH}_2\text{NH}_2\text{Cl} \quad \text{KF} \]
\[ \text{KBr} \quad \text{NH}_4\text{I} \]
\[ \text{KBrO}_3 \quad \text{KClO} \]

Find the pH of a 0.01M \text{CH}_3\text{CH}_2\text{NH}_2 solution.
**pH of Acidic Salts**

In order to determine the pH of acidic salts, it is first important to write the dissociation reaction. Acidic salts contain a conjugate acid. Conjugate acids will be weak acids so most will stay in the $\text{BH}^+$ form. Only a small amount will dissociate into $\text{B}$ and $\text{H}^+$. The calculations are the same as for all weak acids, but you must first determine the $K_a$ by using the formula: $K_a = \frac{1 \times 10^{-14}}{K_b}$

Write the dissociate reaction for $\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl}$. (Remember to get rid of the spectator!):

Determine the $K_a$ for $\text{CH}_3\text{CH}_2\text{NH}_3^+$

Determine the pH for a 0.01 M $\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl}$ solution

**pH of Basic Salts**

In order to determine the pH of basic salts, it is first important to write the dissociation reaction. Basic salts contain a conjugate base. Conjugate bases will be weak bases so most will stay in the $\text{A}^-$ form. Only a small amount will react to form $\text{HA}$ and $\text{OH}^-$. The calculations are the same as for all weak bases, but you must first determine the $K_b$ by using the formula: $K_b = \frac{1 \times 10^{-14}}{K_a}$

Write the reactions for $\text{KNO}_2$. (Remember to get rid of the spectator!):

Determine the $K_b$ for $\text{NO}_2^-$

Determine the pH for a 0.01 M $\text{KNO}_2$ solution

Draw a beaker diagram: